How to balance oxidation-reduction equations using the half-cell method.

Consider the following unbalanced reaction:

$$C + Fe_2O_3 \xrightarrow{\Delta} Fe + CO_2$$

This reaction is paramount in the production of steel.

Let's take a closer look to see what happens:

1st assign oxidation numbers to all atoms of all the elements in the equation

$$C^0 + Fe_2^{+3}O_3^{-2} \xrightarrow{\Delta} Fe^0 + C^{+4}O_2^{-2}$$

Okay, how did I get a +3 for iron? If we look at the negative charge from the 3 oxides the total is -6 ( $3 \times -2 = -6$ ). Since all compound formulas have to total zero (0), the 2 Fe atoms must contribute +6. And each Fe atom will contribute +3. ( $+6 \div 2 = +3$ )

2nd Inspect the equation to see which elements changed oxidation states. In this particular reaction C and Fe change states.

(Hint: 2 elements will change 99.9% of the time)

$C^0 \rightarrow C^{+4}$ and $Fe_2^{+3} \rightarrow 2Fe^0$

Note: the subscript on the left for Fe has become a coefficient on the right side of the half-cell.

3rd Set up the individual half-cells for each element showing how many electrons were lost or gained by the 2 elements

$C^0 - 4e^{-1} \rightarrow C^{+4}$ $Fe_2^{+3} + 6e^{-1} \rightarrow 2Fe^0$

4th Now use coefficients in front of the half-cells where needed to get the number of electrons gained and electrons lost to equal.

a) A common multiple needs to be established to get this to work. Since the number of electrons gained is 6 and the number of electrons lost is 4; 12 seems to work.

b) The half-cells will look like this:

$3(C^0 - 4e^{-1} \rightarrow C^{+4})$ $2(Fe_2^{+3} + 6e^{-1} \rightarrow 2Fe^0)$

c) Treat this like a math equation (the distributive law)

$$2Fe_2^{+3} + 12e^{-1} \rightarrow 4Fe^0$$
$$3C^0 - 12e^{-1} \rightarrow 3C^{+4}$$

Upon inspection you will note that the number of electrons gained (+12) equals the number of electrons lost (-12)

5$^{th}$ Transfer coefficients from completed half-cells to the main equation:

$$3C + 2Fe_2O_3 \rightarrow 4Fe + 3CO_2$$

6$^{th}$ Balance any remaining elements EXCEPT hydrogen and oxygen
(in this case there are none)

7$^{th}$ Balance hydrogen and THEN oxygen.

Upon further review, each side has 6 oxygen atoms.

8$^{th}$ DONE

In this reaction, iron oxide ($Fe_2O_3$) gives away its oxygen to carbon (C). In chemical terms, the carbon is oxidized because it has gained oxygen. At the same time, the iron oxide is reduced because it has lost oxygen.

Because of its ability to give away oxygen, iron oxide is called an oxidizing agent. Similarly, because of its ability to take on oxygen, carbon is said to be a reducing agent. Oxidation and reduction always occur together. If one substance gives away oxygen (oxidation), a second substance must be present to take on that oxygen (reduction).

## Chemistry I: Rules for assigning oxidation numbers

1. An uncombined element (free element) has an oxidation number of zero (0).

2. A monatomic ion (1 atom) has an oxidation number equal to its charge.

3. Fluorine's oxidation number is always –1.

4. Oxygen has an oxidation number of –2 in all compounds except peroxide where the oxidation number is –1.

5. Hydrogen has an oxidation number of +1 except when combined with metals where hydrogen's oxidation number will be –1.
   Ex: NaH (sodium hydride)

6. All Group 1 elements have an oxidation number of +1. All Group 2 elements have an oxidation number of +2.

7. Second element in a binary compound is assigned the oxidation number it would have if it were an ion. (Hint: always negative)

8. The algebraic sum of the oxidation numbers of ALL of the atoms in a compound MUST equal zero.

9. The algebraic sum of the oxidation numbers of ALL the atoms in a polyatomic ion is equal to the charge on the ion. Ex. Sulfate ion is –2 so all the oxidation numbers in this ion must add up to –2.

10. When in doubt ask your chemistry teacher!!!!!!!!!

## Chemistry I: Redox Equations:

**Rules for balancing REDOX equations by half-cell method.**

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| 1. Assign oxidation numbers to each individual element. |
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| 2. Determine which elements change in charge from one side of the equation to the other. Write them below the equation. |
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| 3. Determine how many electrons are lost and gained by each element that changes. These are called half-cells. |
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| 4. Use coefficients to balance the number of electrons gained and lost. The number of electrons lost by one element MUST equal the number of electrons gained by another element. |
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| 5. Use coefficients obtained in the half-cell reactions to balance those elements in the complete equation. |
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| 6. Balance ALL other elements except oxygen and hydrogen. |
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| 7. Balance hydrogen. |
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| 8. Balance oxygen. |
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Hints for identifying the elements that change in a redox reaction.

1. An element is a free element on one side of the equation and combined on the other.
2. Transition metals are involved in the reaction
3. An element changes position
   a. $KMnO_4$ -----> $MnO_2$
   b. $KClO_3$ -----> KCl
4. also sulfates and sulfites can change
   a. $SO_4^{-2}$ -----> $SO_3^{-2}$